Name ____ AP Chemistry

Kinetics of a Reaction

Introduction

How fast will a chemical reaction occur? If a reaction is too slow, it may not be practical. If the reaction is too fast, it may explode. Measuring and controlling reaction rates makes it possible for chemists and engineers to make a variety of products, everything from antibiotics to fertilizers, in a safe and economical manner. The study of the speed of reactions and the processes which occur as the reaction takes place is known as kinetics. The purpose of this experiment is to investigate how the rate of a reaction can be measured and how reaction conditions affect reaction rates.

Background

This experiment is designed to study the kinetics of a chemical reaction. The reaction involves the oxidation of iodide ions by bromate ions in the presence of acid:

$$6\ I^- (aq) + BrO_3^- (aq) + 6\ H^+ (aq) \rightarrow 3\ I_2 (aq) + Br^- (aq) + 3\ H_2O (l)$$

The reaction is somewhat slow at room temperature. The reaction rate depends on the concentration of the reactants and on the temperature. The rate law for the reaction is a mathematical expression that relates the reaction rate to the concentrations of reactants. If the rate of reaction is expressed as the rate of decrease in concentration of bromate ion, the rate law has the form:

$$Rate = -\frac{\Delta[BrO_3^-]}{\Delta t} = k[I^-]^x[BrO_3^-]^y[H^+]^z$$

where the square brackets refer to the molar concentration of the indicated species. The rate is equal to the change in concentration of the bromate ion, $-\Delta[BrO_3^-]$, divided by the change in time for the reaction to occur, Δt. The term “k” is the rate constant for the equation, which changes as the temperature changes. The exponents x, y, and z are called the “orders” of the reaction with respect to the indicated substance, and show how the concentration of each substance affects the rate of reaction. The total rate law for the process is determined by measuring the rate, evaluating the rate constant, k, and determining the order of the reaction for each reactant (the values of x, y, and z).

To find the rate of the reaction a method is needed to measure the rate at which one of the reactants is used up, or the rate at which one of the products is formed. In this experiment, the rate of reaction will be measured based on the rate at which iodine forms. The reaction will be carried out in the presence of thiosulfate ions, which will react with iodine as it forms:

$$I_2\,(aq) + 2\,S_2O_3^{2-}\,(aq) \rightarrow 2\,I^-\,(aq) + S_4O_6^{2-}\,(aq)$$

The first reaction above is somewhat slow. The second reaction is extremely rapid, so that as quickly as iodine is produced in the first reaction, it is consumed in the second reaction. The second reaction continues until all of the added thiosulfate has been used up. After that, iodine begins to increase in concentration in solution. If some starch is present, iodine reacts with the starch to form a deep blue-colored complex that is readily apparent. Carrying out first reaction in the presence of thiosulfate ion and starch produces a chemical "clock." When the thiosulfate is consumed, the solution turns blue almost instantly.

In this laboratory procedure, all of the reactions use the same quantity of thiosulfate ion. The blue color appears when all the thiosulfate is consumed. An examination of the two reactions shows that six moles of $S_2O_3^{2-}$ are needed to react with the three moles of I_2 formed from the reaction of one mole of BrO_3^-. Knowing the amount of thiosulfate used, it is possible to calculate both the amount of I_2 that is formed and the amount of BrO_3^- that has reacted at the time of the color change. The reaction rate is expressed as the decrease in concentration of BrO_3^- ion divided by the time it takes for the blue color to appear.

Catalysts are substances that speed up a reaction, but are not consumed in the reaction. Catalysts work by lowering the overall activation energy of the reaction (an energy barrier that all reactants must surmount for a reaction to take place), thus increasing the rate of the reaction.

The experiment is designed so that the amounts of the reactants that are consumed are small in comparison with the total quantities present. This means that the concentration of reactants is almost unchanged during the reaction, and therefore reaction rate is almost a constant during this time.

Safety Precautions

Dilute hydrochloric acid solution is severely irritating to skin and eyes and is slightly toxic by ingestion and inhalation. Dilute copper (II) nitrate solution is irritating to skin, eyes and mucous membranes and slightly toxic by ingestion. Dilute potassium bromate solution is irritating to a body tissue and slightly toxic by ingestion. Wear chemical splash goggles, chemical-resistant gloves, and a chemical-resistant apron. Wash hands thoroughly with soap and water before leaving the laboratory.

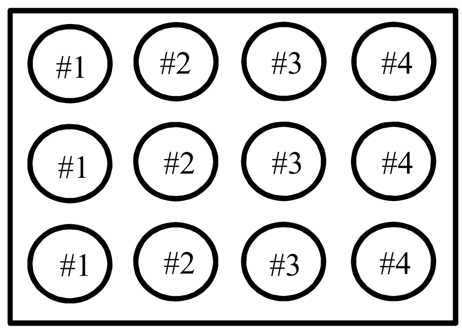
Prelab Questions

1) Explain the relationship between the two reactions described in the background. Why is this known as an "iodine clock"?
2) What is the stoichiometric ratio between $S_2O_3^{2-}$ and BrO_3^-?
3) What is a catalyst and what affect does it have on a reaction? What is unique about catalysts?

Procedure

Part A—Determine the Reaction Rate and Calculate the Rate Law

1) Obtain six pipettes and fold masking tape around the stem of each pipette. Label the pipette: KI, H_2O, HCl, Starch, $Na_2S_2O_3$, and $KBrO_3$.
2) Obtain a clean, 12 well-plate.
3) Using Table #1, put the appropriate number of drops for the first **five (5)** compounds for Experiment #1 into the first well. Repeat for the next two wells below it, as shown below.



4) For the next column of wells, repeat step #3 for Experiment #2.
5) For the next column of wells, repeat step #3 for Experiment #3.
6) For the last column of wells, repeat step #3 for Experiment #4.

Table 1. Reagent Quantities for Experiments

Experiment Number	KI (0.010 M)	Distilled H_2O	HCl (0.10 M)	Starch (2%)	$Na_2S_2O_3$ (0.0010 M)	$KBrO_3$ (0.040 M)
1	2 drops	4 drops	2 drops	1 drop	1 drop	2 drops
2	4 drops	2 drops	2 drops	1 drop	1 drop	2 drops
3	6 drops	0 drops	2 drops	1 drop	1 drop	2 drops
4	2 drops	2 drops	2 drops	1 drop	1 drop	4 drops
5	2 drops	0 drops	2 drops	1 drop	1 drop	6 drops
6	2 drops	2 drops	4 drops	1 drop	1 drop	2 drops
7	2 drops	0 drops	6 drops	1 drop	1 drop	2 drops

7) To well 1, begin a timer as the two drops of $KBrO_3$ are added. Stir the solution with a toothpick.
8) Record in Data Table 1 the time required for the first tint of blue color to appear. (Note: it might be helpful to place the well plate on a piece of white paper to see the color change)
9) Repeat steps #8-9 for the other two wells in that column. Record the times in Data Table 1.
10) Repeat steps #7-9 for Experiments #2-4.
11) Record the temperature of one well for each experiment and record it in Data Table 1.
12) Rinse the well plate with DI water.

13) Setup Experiments #5-7 by adding the first **five (5)** compounds to the each of three wells for each experiment.
14) Repeat steps #7-9 and record all times in Data Table 1. Record the temperature of one well for each experiment in Data Table 1.
15) Rinse the well plate with DI water.

Part B—Observe the Effect of a Catalyst on the Rate

1) Setup Experiment #1 from Table 1 in only one well. Add only the first **five (5)** compounds, except only add 3 drops of DI water. Add one drop of 0.1 M copper (II) nitrate solution.
2) Begin a timer as the two drops of $KBrO_3$ are added. Stir the solution with a toothpick. Record the time in Data Table 2.
3) Rinse the well plate with DI water.

Part C—Find the Volume of One Drop of Solution

1) Obtain a pipette and dry, clean 50 mL beaker.
2) Mass the 50 mL beaker and record the mass in Data Table 3.
3) Fill the pipette with DI water and add five drops of water to the 50 mL beaker. Record the mass in Data Table 3.
4) Add five additional drops and record the new mass for trial 2 in Data Table 3.
5) Repeat a third time and record the final mass in trial 3.
6) Clean up equipment.

Data Table

Data Table 1 –Determine the Reaction Rate and Calculate the Rate Law

	Time (s)				
Experiment Number	**Trial 1**	**Trial 2**	**Trial 3**	**Average**	**Temperature (°C)**
1					
2					
3					
4					
5					
6					
7					

Data Table 2—Observe the Effect of a Catalyst on the Rate

	Reaction Time (s)
Uncatalyzed Reaction	
Catalyzed Reaction	

Data Table 3—Find the Volume of One Drop of Solution

	Mass of empty beaker (a)	g
Trial #1	Mass of beaker + 5 drops of water (b)	g
	Mass of first 5 drops (b-a)	g
	Average mass of 1 drop	g
Trial #2	Mass of beaker + 10 drops of water (c)	g
	Mass of second 5 drops (c-a)	g
	Average mass of 1 drop	g
Trial #3	Mass of beaker + 15 drops of water (d)	g
	Mass of third 5 drops (d-a)	g
	Average mass of 1 drop	g
Average mass of 1 drop (Trials 1-3)		g

Calculations

- Calculate the Volume of One Drop of Solution
  1) Assume the density of water to be 1.00 g/mL. Determine the volume of one drop of water using your data from Data Table 3:

$$Volume\ of\ one\ drop = \frac{mass\ 1\ drop\ (g)}{1.00\ g/mL}\ x\ \frac{1\ L}{1000\ mL}$$

- Calculate the Rate

2) In each reaction there is one drop of 0.0010 M $Na_2S_2O_3$ solution. Using the volume from above and this concentration, calculate the number of moles of $S_2O_3^{2-}$ present in one drop.
3) Using the stoichiometric between $S_2O_3^{2-}$ and BrO_3^-, calculate the moles of BrO_3^- reacted.
4) Determine $[BrO_3^-]$ by dividing the moles of BrO_3^- by the volume of 12 drops (volume of one drop x 12).
5) Calculate the rate of each experiment by dividing $[BrO_3^-]$ by the time of each experiment.

Experiment	**Reaction Rate (M/s)**
1	
2	
3	
4	
5	
6	
7	

- Calculate Initial Concentrations

6) Calculate the initial concentration of each reactant (I^-, BrO_3^-, H^+) for each experiment using the following formulas:

$$[I^-] = \frac{\#\ of\ drops\ x\ 0.010\ M}{12\ drops}$$

$$[BrO_3^-] = \frac{\#\ of\ drops\ x\ 0.040\ M}{12\ drops}$$

$$[H^+] = \frac{\#\ of\ drops\ x\ 0.10\ M}{12\ drops}$$

Initial Concentrations (mol/L)			
Experiment	**$[I^-]$**	**$[BrO_3^-]$**	**$[H^+]$**
1			
2			
3			
4			
5			
6			
7			

- Calculate the Order of Each Reactant
  7) Determine rate order of [I^-] by solving for x using Experiment 1 and 2 (The concentration of BrO_3^- and H^+ are constant in both of the experiments):

$$\frac{Rate\ for\ Experiment\ 1}{Rate\ for\ Experiment\ 2} = \frac{[I^-]^x\ (Experiment\ 1)}{[I^-]^x\ (Experiment\ 2)}$$

  8) Determine rate order of [BrO_3^-] by solving for y using Experiment 1 and 4 (The concentration of I^- and H^+ are constant in both experiments):

$$\frac{Rate\ for\ Experiment\ 1}{Rate\ for\ Experiment\ 4} = \frac{[BrO_3^-]^y\ (Experiment\ 1)}{[BrO_3^-]^y\ (Experiment\ 4)}$$

  9) Determine rate order of [H^+] by solving for z using Experiment 1 and 6 (The concentration of I^- and BrO_3^- are constant in both experimetns):

$$\frac{Rate\ for\ Experiment\ 1}{Rate\ for\ Experiment\ 6} = \frac{[H^+]^z\ (Experiment\ 1)}{[H^+]^z\ (Experiment\ 6)}$$

  10) Write the experimentally determined rate law.

Postlab Questions

1) Why does the reaction rate change as concentrations of the reactants change?
2) Explain the general process used to find the rate law.
3) Describe the effect of a catalyst by comparing the reaction rates.